



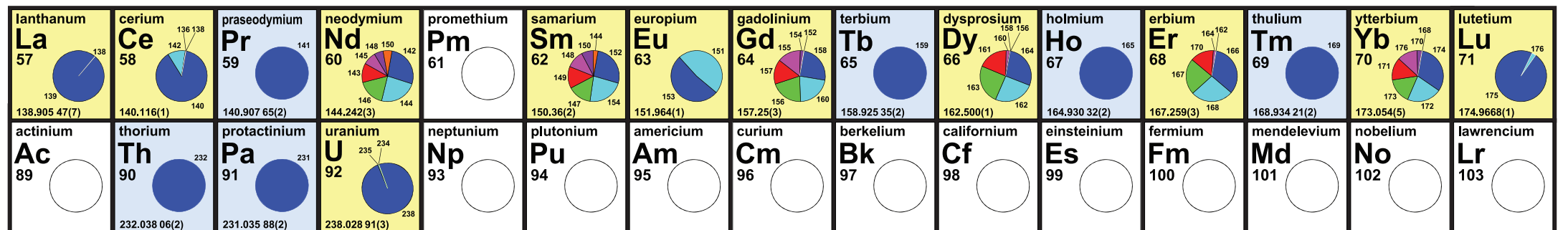
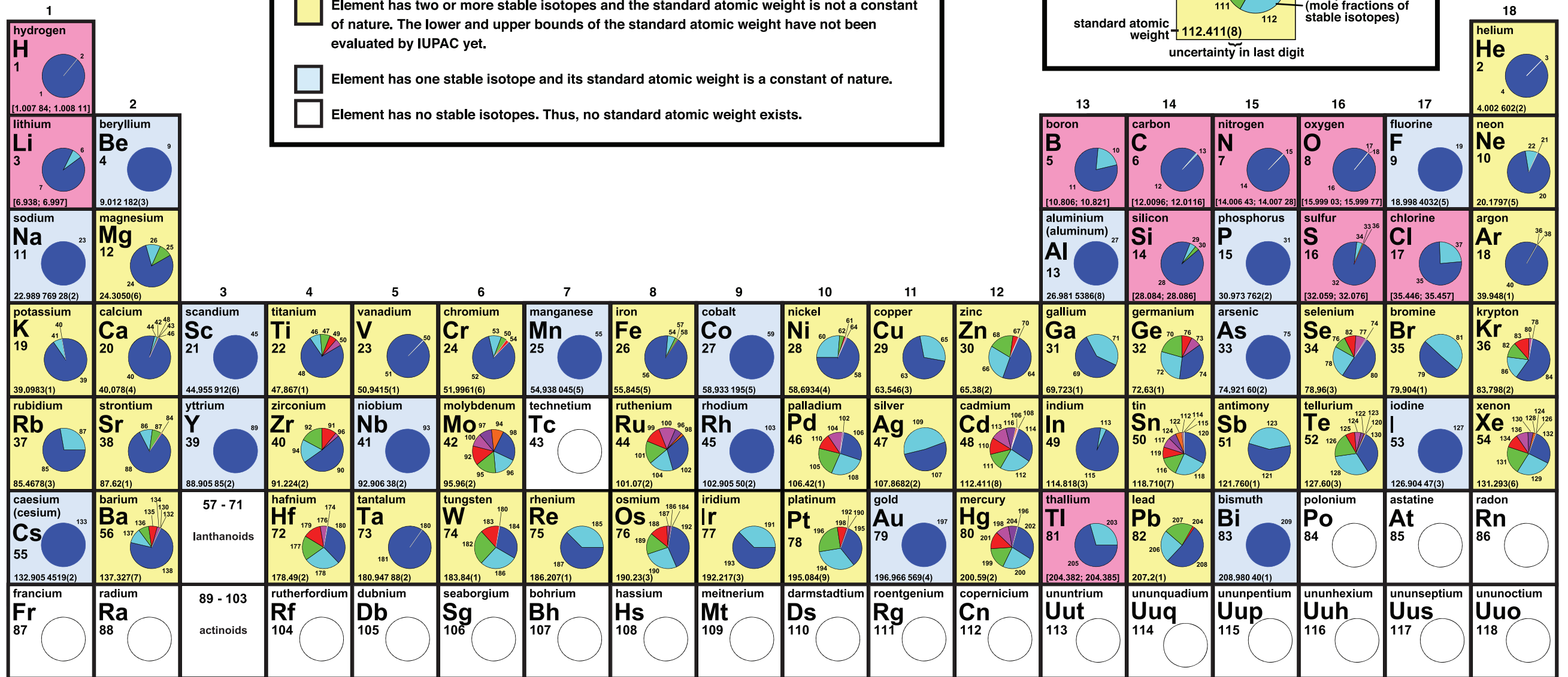
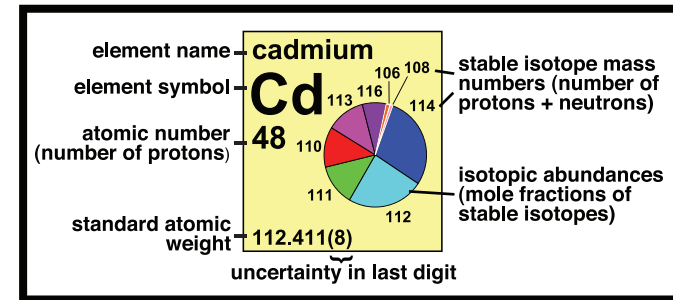


# IUPAC Periodic Table of the Isotopes

-  Element has two or more stable isotopes. Atomic weight and isotopic abundances of element vary in naturally occurring materials. The lower and upper bounds of atomic weight have been assessed by IUPAC and are presented as the standard atomic weight within square brackets, [ ].
-  Element has two or more stable isotopes and the standard atomic weight is not a constant of nature. The lower and upper bounds of the standard atomic weight have not been evaluated by IUPAC yet.
-  Element has one stable isotope and its standard atomic weight is a constant of nature.
-  Element has no stable isotopes. Thus, no standard atomic weight exists.

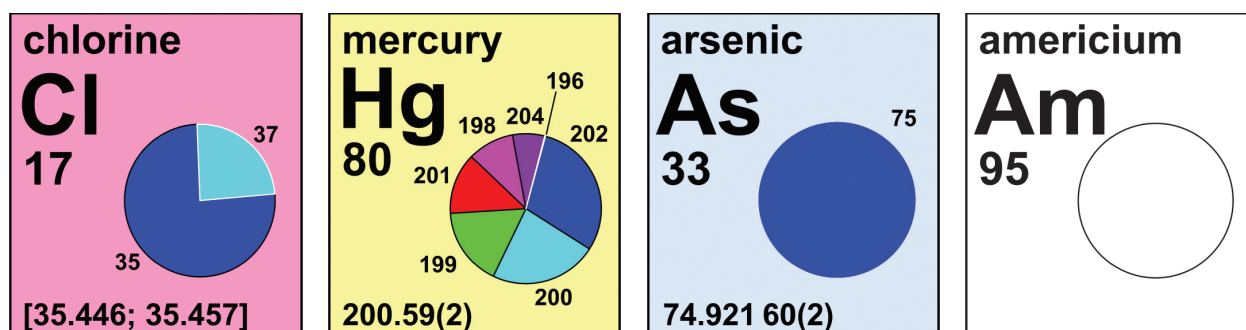


International Year of  
**CHEMISTRY**  
 2011

## IUPAC Periodic Table of the Isotopes

The world that surrounds us, including the water that we drink and the air that we breathe, is mostly made of stable isotopes. The IUPAC Periodic Table of the Isotopes is intended to familiarize students and teachers with the existence of isotopes of chemical elements. The more familiar Periodic Table of the Elements found in most chemistry text books and on classroom walls illustrates important patterns of chemical behavior that are related to systematic variations in electronic configurations and atomic weights. The Periodic Table of the Isotopes highlights several categories of elements that are dif-

ferentiated by the existence or non-existence of stable isotopes and the impact of the variation in stable isotope amounts on the atomic weight. For each element, the number of stable isotopes and their mass numbers are illustrated with colored pie charts; the relative amount of each stable isotope (called the isotopic abundance) is indicated by the size of the pie slice for that mass number. The table shows the current standard atomic weight for each element with its measured uncertainty in the last digit, or in some cases as an upper and lower bound.



Four of the elements in IUPAC's new periodic table of the isotopes for the educational community with stable isotopic abundances shown as pie diagrams (from left to right): Element (chlorine) whose standard atomic weight is not a constant of nature and is an interval. Element (mercury) whose standard atomic weight is not a constant of nature and is not an interval. Element (arsenic) whose standard atomic weight is a constant of nature because it has one stable isotope. Element (americium) that has no stable isotopes and thus no standard atomic weight.

## Isotopes in Everyday Life

A more expanded "interactive" version of the IUPAC Periodic Table of the Isotopes is being developed which highlights applications of isotopes (both stable and unstable) of each chemical element and their practical uses in science and everyday life. Some types of isotope applications include forensics, geochronology, earth-system science, environment and human health science, and medical diagnosis and treatment. Useful unstable or radioactive isotopes (also called radio-isotopes) include technetium-99 ( $^{99}\text{Tc}$ ) for imaging the internal organs of the human body for medical diagnosis; americium-241 ( $^{241}\text{Am}$ )

for activating the alarm in smoke detectors; and carbon-14 ( $^{14}\text{C}$ ) for answering questions such as "how long since the last Ice Age" and "how old is the Shroud of Turin." Useful stable isotopes include boron-10 ( $^{10}\text{B}$ ) used in boron neutron capture therapy for treating brain tumors and carbon-13 and carbon-12 ( $^{13}\text{C}/^{12}\text{C}$ ) or oxygen-18 and oxygen-16 ( $^{18}\text{O}/^{16}\text{O}$ ) to detect performance-enhancing drug use in sports, authenticity of foods, effects of climate change, and origins of contaminants in the environment. Visit [www.ciaaw.org](http://www.ciaaw.org) for more details.

## Isotopic Abundances and Standard Atomic Weights

Each atom of every element has protons and neutrons in its center (called the nucleus):

- The number of protons in each atom (i.e., the **atomic number**, symbol  $Z$ ) determines the *chemical properties* of that atom. All atoms of a given element have the same number of protons (e.g., carbon atoms have 6 protons, while all uranium atoms have 92 protons).
- The number of neutrons (symbol  $N$ ) in an atom of a given element may vary. Atoms of a given element which have different numbers of neutrons are called **isotopes**.
- The total number of protons and neutrons ( $Z + N$ ) in a given isotope is called the **mass number**.

Isotopes may be stable or unstable. The unstable isotopes are also called radio-isotopes and will decay over time into another isotope of the same or a different element. The stable isotopes of each element and their mass numbers appear on each element cell of the Table in a pie chart, where each mass number is indicated around the outside of the pie. The relative amount of each stable isotope in a given element (called the **isotopic abundance**) is approximately indicated by the size of the slice of the pie assigned to each mass number.

Some elements have no stable isotopes and therefore no mass numbers; these elements are indicated in white on the Table. Other elements have only one stable isotope and a single mass number and they are indicated in blue. The remaining elements have more than one stable isotope and mass number; these elements are indicated in either yellow or pink.

The atomic weight of an element is calculated from the sum of the products of the atomic mass<sup>1</sup> and the isotopic abundance<sup>2</sup> of each stable isotope of that element. The **Standard Atomic Weight** is the recommended value that can apply to all samples of a given element. Consider the simplified calculation for the case of carbon (all of whose isotopes have 6 protons since its atomic number is 6). Carbon has 2 stable isotopes of mass number 12 (abbreviated as  $^{12}\text{C}$ ; corresponding to an atom with 6 protons and 6 neutrons) and mass number 13 ( $^{13}\text{C}$ ; corresponding to an atom with 6 protons and 7 neutrons). The atomic mass of each isotope can be approximated by its mass number,  $^{12}\text{C} \approx 12$  and  $^{13}\text{C} \approx 13$ . Natural carbon is a mixture of  $^{12}\text{C}$  and  $^{13}\text{C}$  atoms with approximate isotopic abundances of 99% and 1%, respectively. The approximate atomic weight for this sample of carbon would be  $12 \times 0.99 + 13 \times 0.01 = 12.01$ .

For any element that has two or more stable isotopes, there is always the possibility that the relative amounts of stable isotopes may vary in various samples of that element found in nature. Using the above example, let us assume that another sample of carbon is made up of 98%  $^{12}\text{C}$  and 2%  $^{13}\text{C}$ . This sample of carbon has an approximate atomic weight of  $12 \times 0.98 + 13 \times 0.02 = 12.02$ . Thus, natural isotopic variation for an element can have an effect on the element's atomic weight value. For 10 such elements, the Standard Atomic Weight assigned by IUPAC is given as upper and lower bounds (called an interval)<sup>3,4</sup> written in brackets (e.g., for chlorine, it is [35.446; 35.457]). These elements are indicated in pink on the Table. Those elements for which no such assessments have been made or completed yet are indicated in yellow and their Standard Atomic Weights are given with an uncertainty in parentheses [e.g., for mercury, 200.59(2) is a contracted notation of 200.59 +/- 0.02].

### References

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